CHM129
Acid-Base Equilibrium
Weak Acid – Strong Base and Weak Base – Strong Acid Titrations

1. Consider the titration of 40.0 mL of 0.100M acetic acid ($K_a = 1.8 \times 10^{-5}$) with a 0.200 M NaOH solution.

$$HC_2H_3O_{2\,(aq)} + NaOH_{\,(aq)} \rightarrow H_2O_{\,(l)} + NaC_2H_3O_{2\,(aq)}$$

Determine the pH at the following points:

a. Initial pH (no NaOH added)
b. pH after the addition of 10.00 mL of NaOH
c. pH after the addition of 20.00 mL of NaOH
d. pH after the addition of 30.00 mL of NaOH

Equivalence Point: $C_aV_a = C_bV_b \Rightarrow V_b = \frac{C_aV_a}{C_b} = \frac{(0.100M)(40.00mL)}{(0.200M)} = 20.0mL\ NaOH$

(a) Initial pH

$[H_3O^+] = \sqrt{K_a \times [acid]} = \sqrt{(1.8\times10^{-5})(0.100)} = 1.3\times10^{-3}M$ $\quad pH = -\log[H_3O^+] = -\log(1.3\times10^{-3}) = \underline{2.87}$

(b) Buffer Region

$HC_2H_3O_{2(aq)} + NaOH_{(aq)} \longrightarrow H_2O_{(l)} + NaC_2H_3O_{2(aq)}$

BA	4.00 mmol	2.00 mmol		
A	—	—		—
AA	2.00 mmol	0.00		2.00 mmol

* Note: Midpoint
$V_b = \frac{1}{2}V_{eq} = 10.00mL$
So, $pH = pK_a$

$pH = pK_a + \log\frac{[base]}{[acid]} = -\log(1.8\times10^{-5}) + \log\left(\frac{2.00mmol}{2.00mmol}\right) = \underline{4.74}$ ($\log(1) = 0$)

(c) Equivalence Point

$HC_2H_3O_2 + NaOH \longrightarrow H_2O + NaC_2H_3O_2$

BA	4.00 mmol	4.00 mmol		
A	—	—		—
AA	0.00	0.00		4.00 mmol

$[NaC_2H_3O_2] = \frac{4.00mmol}{60.00mL} = \underline{0.0667M}$

$K_b = \frac{K_w}{K_a} = \frac{1.0\times10^{-14}}{1.8\times10^{-5}} = 5.6\times10^{-10}$

$[OH^-] = \sqrt{K_b \times [base]} = \sqrt{(5.6\times10^{-10})(0.0667)} = 6.1\times10^{-6}M$

$pOH = -\log(6.1\times10^{-6}) = 5.22 \Rightarrow pH = 14.00 - 5.22 = \underline{8.78}$

(d) After Equivalence Point (Strong base excess)

$HC_2H_3O_2 + NaOH \longrightarrow H_2O + NaC_2H_3O_2$

BA	4.00 mmol	6.00 mmol		
A	—	—		—
AA	0.00	2.00 mmol		4.00 mmol

$[OH^-] = [NaOH] = \frac{2.00mmol}{70.00mL} = 0.0286M$

$pOH = -\log(0.0286) = 1.544$

$pH = 14.00 - 1.544 = \underline{12.46}$

2. Consider the titration of 35.0 mL of 0.140M hypobromous acid ($K_a = 2.5 \times 10^{-9}$) with a 0.200 M KOH solution.

$$HBrO_{(aq)} + KOH_{(aq)} \rightarrow H_2O_{(l)} + KBrO_{(aq)}$$

Determine the pH at the following points:

a. Initial pH (no KOH added)
b. pH after the addition of 13.00 mL of KOH
c. pH after the addition of 24.50 mL of KOH
d. pH after the addition of 35.00 mL of KOH

Equivalence Point

$C_aV_a = C_bV_b \Rightarrow V_b = \frac{C_aV_a}{C_b} = \frac{(35.0\,mL)(0.140\,M)}{0.200\,M} = 24.50\,mL$

(a) Initial pH

$[H_3O^+] = \sqrt{K_a \times [acid]} = \sqrt{(0.140M)(2.5\times10^{-9})} = 1.9\times10^{-5}\,M \Rightarrow pH = -\log(1.9\times10^{-5}) = \underline{4.73}$

(b) Buffer Region

	HBrO	+ KOH	→	H_2O +	KBrO
BA	4.90 mmol	—			
A	—	2.60 mmol			2.60 mmol
AA	2.30 mmol	0.00			

$pH = pKa + \log\frac{[base]}{[acid]} = -\log(2.5\times10^{-9}) + \log\frac{(2.60)}{(2.30)} = \underline{8.66}$

(c) Equivalence Point

	HBrO	+ KOH	→	H_2O +	KBrO
BA	4.90 mmol	—			
A	—	4.90 mmol			4.90 mmol
AA	0.00 mmol	0.00			

$[KBrO] = \frac{4.90\,mmol}{59.50\,mL} = 0.0824\,M$

$K_b = \frac{K_w}{K_a} = \frac{1.0\times10^{-14}}{2.5\times10^{-9}} = 4.0\times10^{-6}$

$[OH^-] = \sqrt{K_b \times [KBrO]} = \sqrt{(4.0\times10^{-6})(0.0824)} = 5.7\times10^{-4}\,M$

$pOH = -\log(5.7\times10^{-4}) = 3.24 \Rightarrow pH = 14.00 - 3.24 = \underline{10.76}$

(d) After Equivalence Point (Strong base excess)

	HBrO	+ KOH	→	H_2O +	KBrO
BA	4.90 mmol	—			
A	—	7.00 mmol			4.90 mmol
AA		2.10 mmol			

$[OH^-] = [KOH] = \frac{2.10\,mmol}{70.00\,mL} = 0.0300\,M$

$pOH = -\log(0.0300) = 1.523 \Rightarrow pH = 14.00 - 1.523 = \underline{12.48}$

3. Consider the titration of 30.0 mL of 0.200M ammonia ($K_b = 1.8 \times 10^{-5}$) with a 0.100 M HCl solution.

$$NH_{3\,(aq)} + HCl_{(aq)} \rightarrow NH_4Cl_{(aq)}$$

Determine the pH at the following points:

a. Initial pH (no HCl added)
b. pH after the addition of 30.00 mL of HCl
c. pH after the addition of 60.00 mL of HCl
d. pH after the addition of 70.00 mL of HCl

Equivalence Point:

$C_aV_a = C_bV_b \Rightarrow V_a = \frac{C_bV_b}{C_a} = \frac{(30.0mL)(0.200M)}{(0.100M)} = 60.0\ mL$

(a) Initial pH

$[OH^-] = \sqrt{K_b \times [base]} = \sqrt{(1.8\times10^{-5})(0.200)} = 1.9\times10^{-3}M$

$pOH = -\log(1.9\times10^{-3}) = 2.72 \Rightarrow pH = 14.00 - 2.72 = \underline{11.28}$

(b) Buffer Region

	NH_3	+ HCl	$\rightarrow$	NH_4Cl
BA	6.00 mmol	—		—
A	—	3.00 mmol		—
AA	3.00 mmol	0.00		3.00 mmol

$K_a = \frac{K_w}{K_b} = \frac{1.0\times10^{-14}}{1.8\times10^{-5}}$

$K_a = 5.6\times10^{-10}$

$pH = pK_a + \log\frac{[base]}{[acid]} = -\log(5.6\times10^{-10}) + \log\frac{(3.00)}{(3.00)} = \underline{9.25}$

* Note: Mid point
$V_a = \frac{1}{2}V_{eq} = 30.0\ mL$
$pH = pK_a$

(c) Equivalence Point

	NH_3	+ HCl	$\rightarrow$	NH_4Cl
BA	6.00 mL	—		—
A	—	6.00 mmol		—
AA	0.00	0.00		6.00 mmol

$[NH_4Cl] = \frac{6.00\ mmol}{90.00\ mL} = 0.0667M$

$[H_3O^+] = \sqrt{K_a \times [acid]} = \sqrt{(5.6\times10^{-10})(0.0667)} = 6.1\times10^{-6}M \Rightarrow pH = -\log(6.1\times10^{-6}) = \underline{5.22}$

(d) After Equivalence Point (Strong acid excess)

	NH_3	+ HCl	$\rightarrow$	NH_4Cl
BA	6.00 mmol	—		—
A	—	7.06 mmol		—
AA	0.00	1.00 mmol		6.00 mmol

$[H_3O^+] = [HCl] = \frac{1.00\ mmol}{100.00\ mL} = 0.0100M$

$pH = -\log(0.0100) = \underline{2.000}$